# Chem101:General Chemistry
# Chapter 4: Forces Between Particles

## IV. ANSWERS AND SOLUTIONS TO ODD-NUMBERED PROBLEMS:

### Noble Gas Configurations (Section 4.1)

4.1 Refer to the group numbers of the periodic table and draw Lewis structures for atoms of the following:
a) lithium b) sodium c) chlorine d) boron

**SOLUTION:**

a) Li · Group IA(1) b) Na · Group IA(1) c) $:\ddot{\underset{\cdot\cdot}{\mathrm{Cl}}}\cdot$ Group VIIA(17) d) $\cdot\dot{\mathrm{B}}\cdot$ Group IIIA(13)

4.3 Write abbreviated electronic configurations for the following:
a) iodine b) element number 38 c) As d) phosphorus

**SOLUTION:**
a) [Kr] $5s^2 4d^{10} 5p^5$ b) [Kr] $5s^2$ c) [Ar] $4s^2 3d^{10} 4p^3$ d) [Ne] $3s^2 3p^3$

4.5 Draw Lewis structures for the elements given in Exercise 4.3.

**SOLUTION:**

a) $:\ddot{\underset{\cdot\cdot}{\mathrm{I}}}\cdot$ b) Sr : c) $\cdot\ddot{\underset{\cdot}{\mathrm{As}}}\cdot$ d) $\cdot\ddot{\underset{\cdot}{\mathrm{P}}}\cdot$

4.7 Use the symbol *E* to represent an element in a general way and draw Lewis structures for atoms of the following:
a) any group IA (1) element b) any group IVA (14) element

**SOLUTION:**

a) E · b) $\cdot\dot{\underset{\cdot}{\mathrm{E}}}\cdot$

**Ionic Bonding (Section 4.2)**

4.9 Indicate both the minimum number of electrons that would have to be added and the minimum number that would have to be removed to change the electronic configuration of each element listed in Exercise 4.3 to a noble gas configuration.

**SOLUTION:**
a) add 1 or lose 17 b) add 16 or lose 2 c) add 3 or lose 15 d) add 3 or lose 5

4.11 Use the periodic table and predict the number of electrons that will be lost or gained by the following elements as they change into simple ions. Write an equation using elemental symbols, ionic symbols, and electrons to represent each change.
a) Ca b) aluminum c) fluorine d) element number 34

**SOLUTION:**
a) lose 2; $Ca \rightarrow Ca^{2+} + 2\,e^-$ b) lose 3; $Al \rightarrow Al^{3+} + 3\,e^-$
c) gain 1; $F + e^- \rightarrow F^-$ d) gain 2; $Se + 2\,e^- \rightarrow Se^{2-}$

4.13 Write a symbol for each of the following ions:
a) a bromine atom that has gained one electron
b) a sodium atom that has lost one electron
c) a sulfur atom that has gained two electrons

**SOLUTION:**
a) $Br^-$ b) $Na^+$ c) $S^{2-}$

4.15 Identify the element in period 2 that would form each of the following ions. *E* is used as a general symbol for an element.
a) $E^-$ b) $E^{2+}$ c) $E^{3-}$ d) $E^+$

**SOLUTION:**
a) $F^-$ b) $Be^{2+}$ c) $N^{3-}$ d) $Li^+$

4.17 Identify the noble gas that is isoelectronic with each of the following ions:
a) $Mg^{2+}$ b) $Te^{2-}$ c) $N^{3-}$ d) $Be^{2+}$

**SOLUTION:**
a) Ne b) Xe c) Ne d) He

**Ionic Compounds (Section 4.3)**

4.19 Write equations to represent positive and negative ion formation for the following pairs of elements. Then write a formula for the ionic compound that results when the ions combine.
a) Mg and S b) strontium and nitrogen c) elements number 3 and 34

**SOLUTION:**
a) $Mg \rightarrow Mg^{2+} + 2\,e^-$; $S + 2\,e^- \rightarrow S^{2-}$; MgS
b) $Sr \rightarrow Sr^{2+} + 2\,e^-$; $N + 3\,e^- \rightarrow N^{3-}$; $Sr_3N_2$
c) $Li \rightarrow Li^+ + e^-$; $Se + 2\,e^- \rightarrow Se^{2-}$; $Li_2Se$

4.21 Write the formula for the ionic compound formed from $Sr^{2+}$ and each of the following ions:
a) $S^{2-}$ b) $Br^-$ c) $N^{3-}$ d) $Cl^-$

**SOLUTION:**
a) SrS b) $SrBr_2$ c) $Sr_3N_2$ d) $SrCl_2$

4.23 Classify each of the following as a binary compound or not a binary compound:
a) HF b) $OF_2$ c) $H_2SO_4$ d) $H_2S$ e) $MgBr_2$

**SOLUTION:**
a) binary b) binary c) not binary (ternary) d) binary e) binary

**Naming Binary Ionic Compounds (Section 4.4)**

4.25 Name the following metal ions:
a) $Ca^{2+}$ b) $K^+$ c) $Al^{3+}$ d) $Rb^+$

**SOLUTION:**
a) calcium ion b) potassium ion c) aluminum ion d) rubidium ion

4.27 Name the following nonmetal ions:
a) $Cl^-$ b) $N^{3-}$ c) $S^{2-}$ d) $Se^{2-}$

**SOLUTION:**
a) chloride ion b) nitride ion c) sulfide ion d) selenide ion

4.29 Name the following binary ionic compounds:
a) $K_2O$ b) $SrCl_2$ c) $Al_2O_3$ d) LiBr e) CaS

**SOLUTION:**
a) potassium oxide b) strontium chloride c) aluminum oxide
d) lithium bromide e) calcium sulfide

4.31 Name the following binary ionic compounds using a roman numeral to indicate the charge on the metal ion:
a) $CrCl_2$ and $CrCl_3$ b) CoS and $Co_2S_3$ c) FeO and $Fe_2O_3$ d) $PbCl_2$ and $PbCl_4$

**SOLUTION:**
a) chromium(II) chloride and chromium(III) chloride
b) cobalt(II) sulfide and cobalt(III) sulfide
c) iron(II) oxide and iron(III) oxide
d) lead(II) chloride and lead(IV) chloride

4.33 Name the binary compounds of Exercise 4.31 by adding the endings *-ous* and *-ic* to indicate the lower and higher ionic charges of the metal ion in each pair of compounds. Note: The non-English root for lead (Pb) is *plumb-*.

**SOLUTION:**
a) chromous chloride and chromic chloride
b) cobaltous sulfide and cobaltic sulfide
c) ferrous oxide and ferric oxide
d) plumbous chloride and plumbic chloride

4.35 Write formulas for the following binary ionic compounds:
a) manganese(II) chloride b) iron(III) sulfide c) chromium(II) oxide
d) iron(II) bromide e) tin(II) chloride

**SOLUTION:**
a) $MnCl_2$ b) $Fe_2S_3$ c) CrO d) $FeBr_2$ e) $SnCl_2$

**The Smallest Unit of Ionic Compounds (Section 4.5)**

4.37 Determine the formula weight in atomic mass units for each of the following binary ionic compounds:
a) $Na_2O$ b) FeO c) $PbS_2$ d) $AlCl_3$

**SOLUTION:**
a) FW = 2(22.99) + 16.00 = 61.98 u
b) FW = 55.85 + 16.00 = 71.85 u
c) FW = 207.2 + 2(32.06) = 271.32 (round to tenths) = 271.3 u
d) FW = 26.98 + 3(35.45) = 133.33 u

4.39 Identify the ions that would occupy lattice sites in a solid sample of each compound given in Exercise 4.37.

**SOLUTION:**
a) two $Na^+$ ions and one $O^{2-}$ ion
b) one $Fe^{2+}$ ion and one $O^{2-}$ion
c) one $Pb^{4+}$ ion and two $S^{2-}$ ions
c) one $Al^{3+}$ion and three $Cl^-$ ions

4.41 Calculate the mass in grams of positive ions and negative ions contained in 1 mol of each compound given in Exercise 4.37.

**SOLUTION:**
a) One mol $Na_2O$ contains 2 mol $Na^+$ and 1 mol $O^{2-}$. The $Na^+$ ion and $O^{2-}$ ion have the same mass as an Na atom and an O atom.

$$2\ \cancel{\text{mol Na}^+} \times \frac{22.99\text{ g Na}^+}{1\ \cancel{\text{mol Na}^+}} = 45.98\text{ g Na}^+$$

$$1\ \cancel{\text{mol O}^{2-}} \times \frac{16.00\text{ g O}^{2-}}{1\ \cancel{\text{mol O}^{2-}}} = 16.00\text{ g O}^{2-}$$

b) One mol FeO contains 1 mol $Fe^{2+}$ and 1 mol $O^{2-}$. The $Fe^{2+}$ ion and $O^{2-}$ ion have the same mass as an Fe atom and an O atom.

$$1\ \cancel{\text{mol Fe}^{2+}} \times \frac{55.85\text{ g Fe}^{2+}}{1\ \cancel{\text{mol Fe}^{2+}}} = 55.85\text{ g Fe}^{2+}$$

$$1\ \cancel{\text{mol O}^{2-}} \times \frac{16.00\text{ g O}^{2-}}{1\ \cancel{\text{mol O}^{2-}}} = 16.00\text{ g O}^{2-}$$

c) One mol $PbS_2$ contains 1 mol $Pb^{4+}$ and 2 mol $S^{2-}$. The $Pb^{4+}$ ion and $S^{2-}$ ion have the same mass as an Pb atom and a S atom.

$$1\ \cancel{\text{mol Pb}^{4+}} \times \frac{207.2\text{ g Pb}^{4+}}{1\ \cancel{\text{mol Pb}^{4+}}} = 207.2\text{ g Pb}^{4+}$$

$$2\ \cancel{\text{mol S}^{2-}} \times \frac{32.06\text{ g S}^{2-}}{1\ \cancel{\text{mol S}^{2-}}} = 64.12\text{ g S}^{2-}$$

d) One mol $AlCl_3$ contains 1 mol $Al^{3+}$ and 3 mol $Cl^-$. The $Al^{3+}$ ion and $Cl^-$ ion have the same mass as an Al atom and a Cl atom.

$$1\ \cancel{\text{mol Al}^{3+}} \times \frac{26.98\text{ g Al}^{3+}}{1\ \cancel{\text{mol Al}^{3+}}} = 26.98\text{ g Al}^{3+}$$

$$3\ \cancel{\text{mol Cl}^-} \times \frac{35.45\ \text{g Cl}^-}{1\ \cancel{\text{mol Cl}^-}} = 106.35\ \text{g Cl}^-$$

4.43 Calculate the number of positive ions and negative ions contained in 1.00 mol of each compound given in Exercise 4.37.

**SOLUTION:**

Using the results of Exercise 4.39 and knowing that 1 mol = $6.02 \times 10^{23}$ objects, we get:

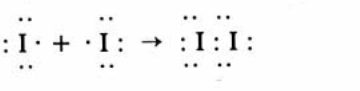
a) $12.0 \times 10^{23}$ $Na^+$ ions and $6.02 \times 10^{23}$ $O^{2-}$ ions
b) $6.02 \times 10^{23}$ $Fe^{2+}$ ions and $6.02 \times 10^{23}$ $O^{2-}$ ions
c) $6.02 \times 10^{23}$ $Pb^{4+}$ ions and $12.0 \times 10^{23}$ $S^{2-}$ ions
d) $6.02 \times 10^{23}$ $Al^{3+}$ ions and $18.1 \times 10^{23}$ $Cl^-$ ions

## Covalent Bonding (Section 4.6)

4.45 Represent the following reaction using Lewis structures:

$$I + I \rightarrow I_2$$

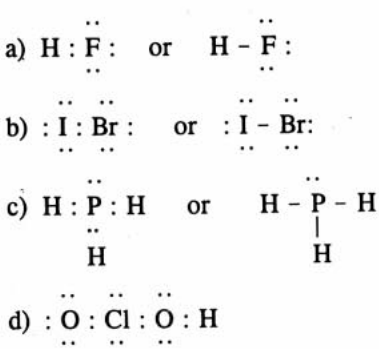
**SOLUTION:**



4.47 Represent the following molecules by Lewis structures:

a) HF b) IBr c) $PH_3$ (each H atom is bonded to the P atom)
d) $HClO_2$ (the O atoms are each bonded to the Cl, and the H is bonded to one of the O atoms)

**SOLUTION:**



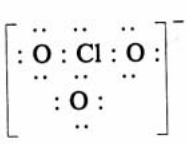
## Polyatomic Ions (Section 4.7)

4.49 Draw Lewis structures for the following polyatomic ions:

a) $ClO_3^-$ (each O atom is bonded to the Cl atom)
b) $CN^-$
c) $CO_3^{2-}$ (each O atom is bonded to the C atom)

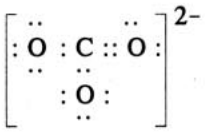
**SOLUTION:**

a) 26 electrons



b) 10 electrons

c) 24 electrons



Note: The double bond could be between any oxygen atom and the carbon atom.

## Shapes of Molecules and Polyatomic Ions (Section 4.8)

4.51 Draw Lewis structures for the following molecules:

a) $O_3$ (the O atoms are bonded together, like beads on a string)
b) $CS_2$ (each S atom is bonded to the C atom)
c) $SeO_2$ (each O atom is bonded to the Se atom)
d) $H_2SO_3$ (each O atom is bonded to the S atom and one H atom is bonded to each of two O atoms)

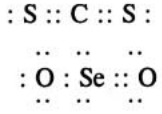
**SOLUTION:**

a) 18 electrons

Note: The double bond and the single bond could be reversed.

b) 16 electrons

c) 18 electrons



Note: The double bond and the single bond could be reversed.

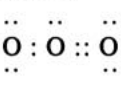
d) 26 electrons

4.53 Predict the shape of each of the following molecules by first drawing a Lewis structure, then applying the VSEPR theory:

a) $O_3$
b) $SeO_2$
c) $PH_3$ (each H atom is bonded to the P atom)
d) $SO_3$ (each O atom is bonded to the S atom)

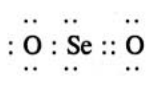
**SOLUTION:**

a)



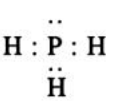
The central oxygen atom has three groups around it, a lone pair and two other oxygen atoms. There will be a triangular (trigonal) arrangement for the orbitals, the three atoms form an angular shape.

b) 18 electrons



The central selenium atom has three groups around it, two lone pairs and two bonded to oxygen atoms. There will be a tetrahedral arrangement for the four orbitals, with the atoms forming an angular shape.

c)



There are four pairs of electrons tetrahedrally arranged around the central P atom. The P and the three H atoms form a triangular pyramid.

d)

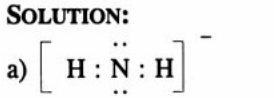


There are three groups trigonally arranged around the S atom. The shape of the molecule is trigonal planar.

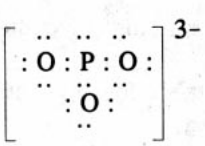
4.55 Predict the shape of each of the following polyatomic ions by first drawing a Lewis structure, then applying the VSEPR theory:

a) $NH_2^-$ ( each H atom is bonded to the N atom)
b) $PO_3^{3-}$ (each O atom is bonded to the P atom)
c) $BeCl_4^{2-}$ (each Cl atom is bonded to the Be atom)
d) $ClO_4^-$ (each O atom is bonded to the Cl atom)

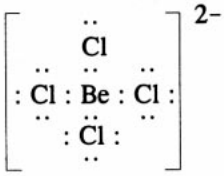
**SOLUTION:**

a) 

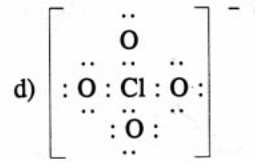
There are four groups, two lone pairs and two H atoms, tetrahedrally arranged around the N atom. The shape of the ion is angular.

b) 

There are four groups, one lone pair and three O atoms, tetrahedrally arranged around the P atom. The shape of the ion is a triangular pyramid.

c) 

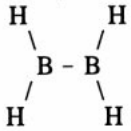
There are four groups tetrahedrally arranged around the Be atom. The shape of the ion is tetrahedral.

d) 

There are four groups tetrahedrally arranged around the Cl atom. The shape of the ion is tetrahedral.

**Polarity of Covalent Molecules (Section 4.9)**

4.57 Use the periodic table and Table 4.3 to determine which of the following bonds will be polarized. Show the resulting charge distribution in those molecules that contain polarized bonds.

a) Cl – F  b) H – Se(–H)  c) 

**SOLUTION:**

a) polar $^{\delta-}Cl - F^{\delta+}$  b) both bonds polar $^{\delta+}H - Se^{\delta-}$
c) B – B bond is non-polar; H – B bonds are polar $^{\delta-}H - B^{\delta+}$

4.59 Use table 4.3 and classify the bonds in the following compounds as *non-polar covalent, polar covalent,* or *ionic.*

a) $PCl_3$  b) $H_2Se$  c) $MgCl_2$  d) $BeI_2$  e) $NCl_3$

**SOLUTION:**

a) polar covalent $\Delta EN = 0.9$  b) polar covalent $\Delta EN = 0.3$  c) polar covalent $\Delta EN = 1.8$
d) polar covalent $\Delta EN = 1.0$  e) nonpolar covalent $\Delta EN = 0$

4.61 On the basis of the charge distributions you drew for the molecules of Exercise 4.57, classify the molecules as *polar* or *nonpolar*.

**SOLUTION:**

a) polar; the bonds are not symmetrical in ClF.
b) polar; the bonds are not symmetrical in $H_2Se$.
c) non-polar; The polar bonds are symmetrical in $B_2H_4$

4.63 Use Table 4.3 and predict the type of bond you would expect to find in compounds formed from the following elements:
a) sulfur and oxygen b) aluminum and bromine c) C and Cl

**SOLUTION:**
a) polar covalent $\Delta EN = 1.0$ b) polar covalent $\Delta EN = 1.3$ c) polar covalent $\Delta EN = 1.5$

4.65 Show the charge distribution in the following molecules and predict which are polar molecules.
a) S = C = S b) H – C ≡ N c) F – O
\
F

**SOLUTION:**
a) nonpolar; the polar bonds are symmetrical $^{\delta-}S = C^{\delta+} = S^{\delta-}$ (linear)
b) polar; $^{\delta+}H - C \equiv N^{\delta-}$ (linear)
c) polar; $^{\delta-}F - O^{\delta+} - F^{\delta-}$ (angular)

**More about Naming Compounds (Section 4.10)**

4.67 Name the following binary covalent compounds:
a) $SiO_2$ b) $SiF_4$ c) $P_2O_5$ d) $AlBr_3$ e) $CBr_4$

**SOLUTION:**
a) silicon dioxide b) silicon tetrafluoride c) phosphorus pentoxide
d) aluminum bromide e) carbon tetrabromide

4.69 Write formulas for the following binary covalent compounds:
a) disulfur monoxide b) sulfur hexafluoride c) silicon tetrachloride d) carbon diselenide

**SOLUTION:**
a) $S_2O$ b) $SF_6$ c) $SiCl_4$ d) $CSe_2$

4.71 Write the formulas and names for compounds composed of ions of the following metals and the indicated polyatomic ions.
a) calcium and the phosphate ion b) sodium and the dichromate ion c) Li and $CO_3^{2-}$ d) Na and $PO_4^{3-}$

**SOLUTION:**
a) $Ca_3(PO_4)_2$ calcium phosphate b) $Na_2Cr_2O_7$ sodium dichromate
c) $Li_2CO_3$ lithium carbonate d) $Na_3PO_4$ sodium phosphate

4.73 Write formulas for the compounds named below:
a) potassium permanganate b) calcium hydroxide c) calcium phosphate
d) ammonium dihydrogenphosphate e) calcium hypochlorite

**SOLUTION:**
a) $KMnO_4$ b) $Ca(OH)_2$ c) $Ca_3(PO_4)_2$ d) $NH_4H_2PO_4$ e) $Ca(ClO)_2$

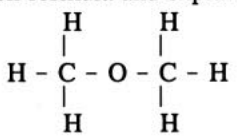
4.75 Write a formula for the following compounds, using M with appropriate charges to represent the metal ion:

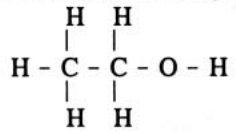
a) any group IIA(2) element and $HSO_3^-$
b) any group IIA(2) element and $HPO_4^{2-}$
c) any metal that forms $M^+$ ions and $NO_2^-$
d) any metal that forms $M^{3+}$ ions and $CO_3^{2-}$
e) any metal that forms $M^{2+}$ ions and $HPO_4^{2-}$

**SOLUTION:**
a) $M(HSO_3)_2$ b) $MHPO_4$ c) $MNO_2$ d) $M_2(CO_3)_3$ e) $MHPO_4$

**Other Interparticle Forces (Section 4.11)**

4.77 The following structural formulas represent molecules of ethyl alcohol and dimethyl ether. Assign the correct name to each formula and explain how your choice is consistent with your answer to Exercise 4.29.





**SOLUTION:**
The structure on the left is nonpolar because of the symmetry; hence, it must be the lower boiling compound, dimethyl ether. The one on the right is polar and must be the higher boiling compound, ethyl alcohol.

4.79 Use the concept of interparticle forces to propose an explanation for the fact that $CO_2$ is a soft, low-melting solid (dry ice), whereas $SiO_2$ is a hard solid (sand). Focus on the nature of the particles that occupy lattice sites in the solid.

**SOLUTION:**
In solid $CO_2$, the lattice site is occupied by a molecule of $CO_2$. The attractive forces are the dispersion forces associated with the nonpolar $CO_2$. In solid $SiO_2$, the lattice sites are occupied by oxygen, silicon, and oxygen, covalently bonded between lattice sites.

4.81 The formula for sucrose is $C_{12}H_{22}O_{11}$, where many of the hydrogen and oxygens are combined to form OH groups that are bonded to carbon atoms. What type of predominant interparticle bonding would you now propose for solid sucrose (see Exercise 4.80.)

**SOLUTION:**
The –OH groups will have polar bonds capable of forming hydrogen bonds between the molecules.